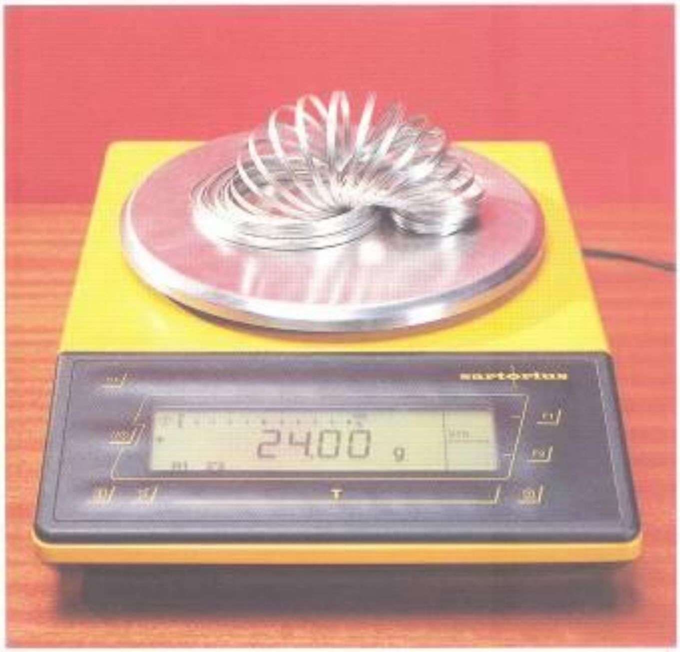
Chapter 5 – Chemical Calculations

Chemists often need to know how much of a sub-stance has been formed or used up during a chemical reaction (Figure 5.1). This is particularly important in the chemical industry, where the substances being reacted (the **reactants**) and the substances being produced (the **products**) are worth thousands of pounds. Waste costs money!

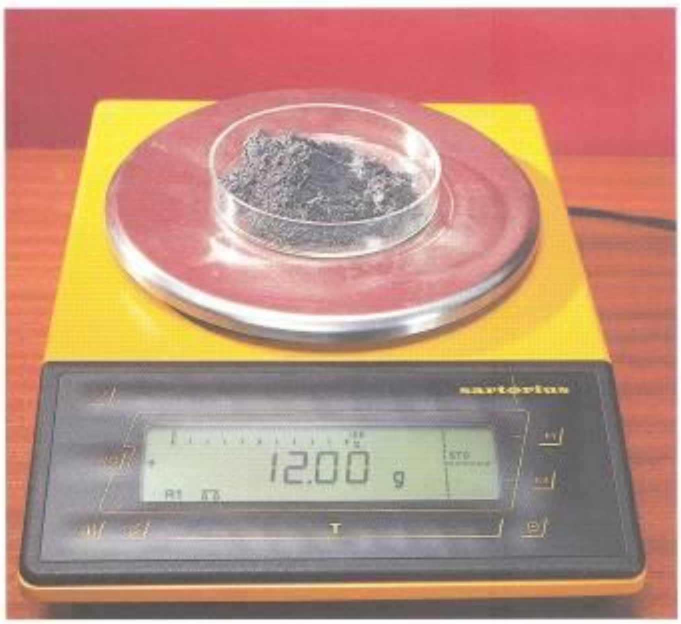
To solve this problem they need a way of counting atoms, ions or molecules. Atoms, ions and molecules are very tiny particles and it is impossible to measure out a dozen or even a hundred of then. Instead, chemists weigh out a very large number of particles. This number is 6×10^{23} atoms, ions or molecules and is called **Avogadro's constant** after the famous Italian scientist Amedeo Avogadro (1776-1856). An amount of substance containing 6×10^{23} particles is called a **mole** (often abbreviated to mol).

So, a mole of the element magnesium is 6×10^{23} atoms of magnesium and a mole of the element carbon is 6×10^{23} atoms of carbon (Figure 5.2).

Figure 5.1 The chemists at paint manufacturers need to know how much pigment is going to be produced.



a A mole of magnesium.



b A mole of carbon.

Figure 5.2

Calculating moles

You have already seen in Chapter 3 how we can compare the masses of all the other atoms with the mass of carbon atoms. This is the basis of the **relative atomic mass scale**. Chemists have found by experiment that if you take the relative atomic mass of an element in grams, it always contains 6×10^{23} or one mole of its atoms.

Moles and elements

For example, the relative atomic mass (A_r) of iron is 56, so one mole of iron is 56 g. Therefore, 56 g of iron contains 6×10^{23} atoms.

The A_r for aluminium is 27. In 27 g of aluminium it is found that there are 6×10^{23} atoms. Therefore, 27 g of aluminium is one mole of aluminium atoms.

The mass of a substance present in any number of moles can be calculated using the relationship:

mass (in grams) = number of moles × mass of 1 mole of the element

Example 1

Calculate the mass of (a) 2 moles and (b) 0.25 mole of iron (A_r: Fe = 56).

a. mass of 2 moles of iron
= number of moles × relative atomic mass (A_r)
= 2 × 56
= 112 g

b. mass of 0.25 mole of iron
= number of moles × relative atomic mass (A_r)
= 0.25 × 56
= 14 g

If we know the mass of the element then it is possible to calculate the number of moles of that element using:

$$\text{number of moles} = \frac{\text{mass of the element}}{\text{mass of 1 mole of that element}}$$

Example 2

Calculate the number of moles of aluminium present in (a) 108 g and (b) 13.5 g of the element (A_r: Al= 27).

a. number of moles of aluminium

$$= \frac{\text{mass of aluminium}}{\text{mass of 1 mole of aluminium}}$$

$$= \frac{108}{27}$$

= 4 moles

b. number of moles of aluminium

$$= \frac{\text{mass of aluminium}}{\text{mass of 1 mole of aluminium}}$$

$$= \frac{13.5}{27}$$

= 0.5 moles

Moles and compounds

The idea of the mole has been used so far only with elements and atoms. However, it can also be used with compounds (Figure 5.3).

We cannot discuss the atomic mass of a molecule or of a compound because more than one type of atom is involved. Instead, we have to discuss the **relative formula mass (RFM)**. This is the sum of the relative atomic masses of all those elements shown in the formula of the substance.

What is the mass of 1 mole of water (H_2O) molecules? (A_r: H = 1; O = 16)

From the formula of water, H_2O, you will see that 1 mole of water molecules contains 2 moles of hydrogen (H) atoms and 1 mole of oxygen (O) atoms. The mass of 1 mole of water molecules is therefore:

$$(2 \times 1) + (1 \times 16) = 18 \text{ g}$$

The mass of 1 mole of a compound is called its molar mass. If you write the molar mass of a compound without any units then it is the relative formula mass, often called the **relative molecular mass (M_r)**. So the relative formula mass of water is 18.

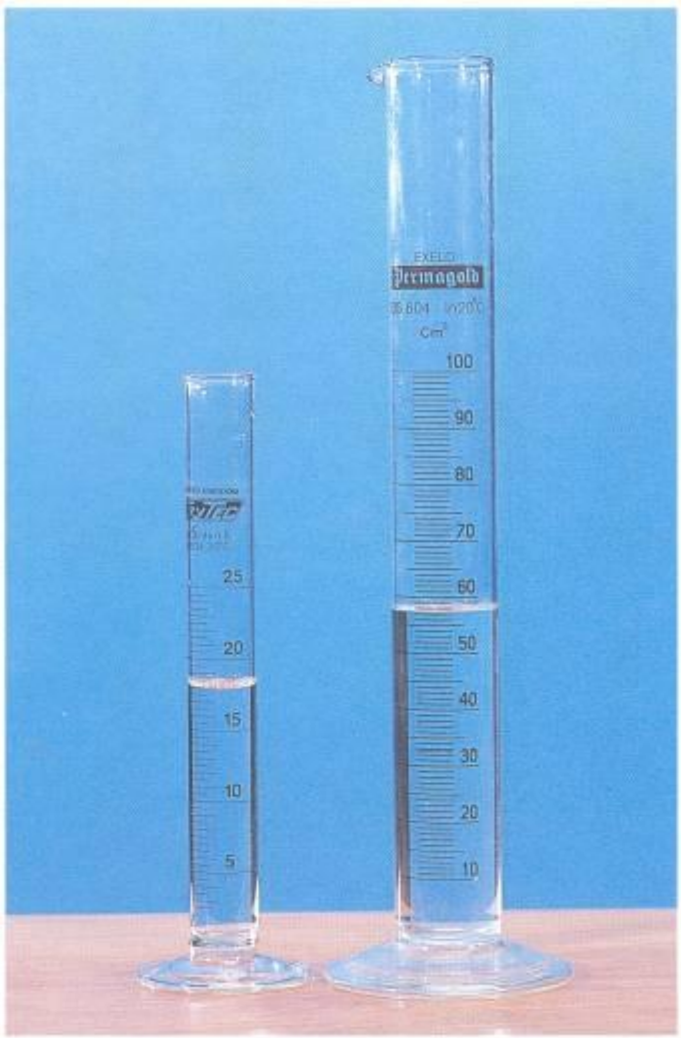


Figure 5.3 One mole of water (H_2O) (left) and one mole of ethanol (C_2H_5OH) (right) in separate measuring cylinders.

Now follow these examples to help you learn and understand more about moles and compounds.

Example 1

What is (a) the mass of 1 mole and (b) the relative formula mass (RFM) of ethanol, C_2H_5OH? (A_r: H = 1; C = 12; O =16)

a. One mole of C_2H_5OH contains 2 moles of carbon atoms, 6 moles of hydrogen atoms and 1 mole of oxygen atoms. Therefore:

mass of 1 mole of ethanol

$= (2 \times 12) + (6 \times 1) + (1 \times 16)$

= 46g

b. The REM of ethanol is 46.

Example 2

What is (a) the mass of 1 mole and (b) the RFM of nitrogen gas, N_2? (A_r: N = 14)

a. Nitrogen is a diatomic gas. Each nitrogen molecule contains two atoms of nitrogen. Therefore:

mass of 1 mole of N_2

$= 2 \times 14$

= 28 g

b. The RFM of N_2 is 28.

The mass of a compound found in any number of moles can be calculated using the relationship:

$$\text{mass of compound} = \text{number of moles of the compound} = \text{mass of 1 mole of the compound}$$

Example 3

Calculate the mass of (a) 3 moles and (b) 0.2 moles of carbon dioxide gas, CO_2. (A_r: C = 12; O = 16)

a. One mole of CO_2 contains 1 mole of carbon atoms and 2 moles of oxygen atoms. Therefore:

mass of 1 mole of CO_2

= (1 × 12) + (2 × 16)

= 44 g

mass of 3 moles of CO_2

= number of moles × mass of 1 mole of CO_2

= 3 × 44

= 132 g

b. mass of 0.2 mole of CO_2

= number of moles × mass of 1 mole of CO_2

= 0.2 × 44

= 8.8 g

If we know the mass of the compound then we can calculate the number of moles of the compound using the relationship:

$$\text{number of moles of the compound} = \frac{\text{mass of compound}}{\text{mass of 1 mole of the compound}}$$

Example 4

Calculate the number of moles of magnesium oxide, MgO, in (a) 80 g and (b) 10 g of the compound. (A_r: O = 16; Mg = 24)

a. One mole of MgO contains 1 mole of magnesium atoms and 1 mole of oxygen atoms. Therefore:

mass of 1 mole of MgO

$= (1 \times 24) + (1 \times 16)$

= 40 g

number of moles of MgO in 80 g

$$= \frac{\text{mass of MgO}}{\text{mass of 1 mole of MgO}}$$

$$= \frac{80}{40}$$

= 2 moles

b. number of moles of MgO in 10 g

$$= \frac{\text{mass of MgO}}{\text{mass of 1 mole of MgO}}$$

$$= \frac{10}{40}$$

= 0.25 mole

Moles and gases

Many substances exist as gases. If we want to find the number of moles of a gas we can do this by measuring the volume rather than the mass.

Chemists have shown by experiment that:

One mole of any gas occupies a volume of approximately 24 dm^3 (24 liter) at room temperature and pressure (rtp).

Therefore, it is relatively easy to convert volumes of gases into moles and moles of gases into volumes using the following relationship:

$$\text{number of moles of a gas} = \frac{\text{volume of the gas (in } dm^3 \text{ at rtp)}}{24\ dm^3}$$

or

$$\text{volume of a gas (in } dm^3 \text{ at rtp)} = \text{number of moles of gas} \times 24 dm^3$$

Example 1

Calculate the number of moles of ammonia gas, NH_3, in a volume of 72 dm^3 of the gas measured at rtp.

$$\text{number of moles of ammonia} = \frac{\text{volume of ammonia in } dm^3}{24\ dm^3}$$

$$= \frac{72}{24}$$

$$= 3$$

Example 2

Calculate the volume of carbon dioxide gas, CO_2, occupied by (a) 5 moles and (b) 0.5 mole of the gas measured at rtp.

a. volume of CO_2

$= \text{number of moles of } CO_2 \times 24\ dm^3$

$= 5 \times 24$

$= 120\ dm^3$

b. volume of CO_2

$= \text{number of moles of } CO_2 \times 24\ dm^3$

$= 0.5 \times 24$

$= 12\ dm^3$

The volume occupied by one mole of any gas must contain 6×10^{23} molecules. Therefore, it follows that equal volumes of all gases measured at the same temperature and pressure must contain the same number of molecules. This idea was also first put forward by Amedeo Avogadro and is called **Avogadro's Law**.

Moles and solutions

Chemists often need to know the concentration of a solution. Sometimes it is measured in grams per cubic decimetre ($g{\cdot}dm^{-3}$) but more often concentration is measured in **moles per cubic decimetre ($mol{\cdot}dm^{-3}$)**. When 1 mole of a substance is dissolved in water and the solution is made up to 1 dm^3 (1000 cm^3), a **1 molar (1 M)** solution is produced. Chemists do not always need to make up such large volumes of solution. A simple method of calculating the concentration is by using the relationship:

$$\text{concentration (in } mol{\cdot}dm^{-3}) = \frac{\text{number of moles}}{\text{volume (in } dm^3)}$$

Example 1

Calculate the concentration (in $mol{\cdot}dm^{-3}$) of a solution of sodium hydroxide, NaOH, which was made by dissolving 10 g of solid sodium hydroxide in

water and making up to 250 cm^3. (A_r: Na = 23; O = 16; H = 1)

1 mole of NaOH contains 1 mole of sodium, 1 mole of oxygen and 1 mole of hydrogen. Therefore:

mass of 1 mole of NaOH = (1 × 23) + (1 × 16) + (1 × 1) = 40 g

number of moles of NaOH in 10 g

$$= \frac{\text{mass of NaOH}}{\text{mass of 1 mole of NaOH}}$$

$$= \frac{10}{40}$$

= 0.25

$$(250\ cm^3 = \frac{250}{1000}\ dm^3 = 0.25\ dm^3)$$

concentration of the NaOH solution

$$= \frac{\text{number of moles of NaOH}}{\text{volume of solution}\ (dm^3)}$$

$$= \frac{0.25}{0.25}$$

= 1 $mol \cdot dm^{-3}$ (or 1 M)

Sometimes chemists need to know the mass of a sub-stance that has to be dissolved to prepare a known volume of solution at a given concentration. A simple method of calculating the number of moles and so the mass of substance needed is by using the relationship:

$$\begin{matrix}\text{number of} \\ \text{moles}\end{matrix} = \begin{matrix}\text{concentration} \\ (\text{in } mol \cdot dm^{-3})\end{matrix} = \begin{matrix}\text{volume of solution} \\ (\text{in } dm^3)\end{matrix}$$

Example 2

Calculate the mass of potassium hydroxide, KOH, that needs to be used to prepare 500 cm^3 of a 2 $mol \cdot dm^{-3}$ (2 M) solution in water. (A_r: H = 1; O = 16; K = 39)

number of moles of KOH

$$= \frac{\text{concentration of solution}}{(\text{mol} \cdot \text{dm}^{-3})} \times \frac{\text{volume of solution}}{(\text{in dm}^3)}$$

$$= 2 \times \frac{500}{1000}$$

= 1

1 mole of KOH contains 1 mole of potassium, 1 mole of oxygen and 1 mole of hydrogen. Therefore:

mass of 1 mole of KOH = (1 × 39) + (1 × 16) + (1 × 1)

= 56 g

Therefore:

mass of KOH in 1 mole

= number of moles × mass of 1 mole

= 1 × 56

= 56 g

Questions

Use the values of A_r which follow to answer the questions below:

H = 1; C = 12; N = 14; O = 16; Ne = 20; Na = 23; Mg = 24; S = 32; K = 39; Fe = 56; Cu = 63.5; Zn = 65.

One mole of any gas at rtp occupies 24 dm^3.

1. Calculate the number of moles in:

a. 2 g of neon atoms

b. 4 g of magnesium atoms

c. 24 g of carbon atoms.

2. Calculate the mass of:

a. 0.1 mole of oxygen molecules

b. 5 moles of sulphur atoms

c. 0.25 mole of sodium atoms.

3. Calculate the number of moles in:

a. 9.8 g of sulphuric acid (H_2SO_4)

b. 40 g of sodium hydroxide (NaOH)

c. 720 g of iron (II) oxide (FeO).

4. Calculate the mass of:

a. 2 moles of zinc oxide (ZnO)

b. 0.25 mole of hydrogen sulphide (H_2S)

c. 0.35 mole of copper (II) sulphate ($CuSO_4$).

5. Calculate the number of moles at rtp in:

a. 2 dm^3 of carbon dioxide (CO_2)

b. 240 dm^3 of sulphur dioxide (SO_2)

c. 20 cm^3 of carbon monoxide (CO).

6. Calculate the volume of:

a. 0.3 mole of hydrogen chloride (HCl)

b. 4.4 g of carbon dioxide

c. 34 g of ammonia (NH_3)

7. Calculate the concentration of solutions containing:

a. 0.2 mole of sodium hydroxide dissolved in water and made up to 100 cm^3

b. 9.8 g of sulphuric acid dissolved in water and made up to 500 cm^3.

8. Calculate the mass of:

a. copper (II) sulphate ($CuSO_4$) which needs to be used to prepare 500 cm^3 of a 0.1 $mol \cdot dm^{-3}$ solution

b. potassium nitrate (KNO_3) which needs to be used to prepare 200 cm^3 of a 2 $mol \cdot dm^{-3}$ solution.

Calculating formulae

If we have 1 mole of a compound, then the formula shows the number of moles of each element in that compound. For example, the formula for lead (II) bromide is $PbBr_2$. This means that 1 mole of lead (II) bromide contains 1 mole of lead ions and 2 moles of bromide ions. If we do not know the formula of a compound, we can find the masses of the elements present experimentally and these masses can be used to work out the formula of that compound.

Finding the formula

Magnesium oxide

When magnesium ribbon is heated strongly, it burns very brightly to form the white powder called magnesium oxide.

$$\text{magnesium} + \text{oxygen} \rightarrow \text{magnesium oxide}$$

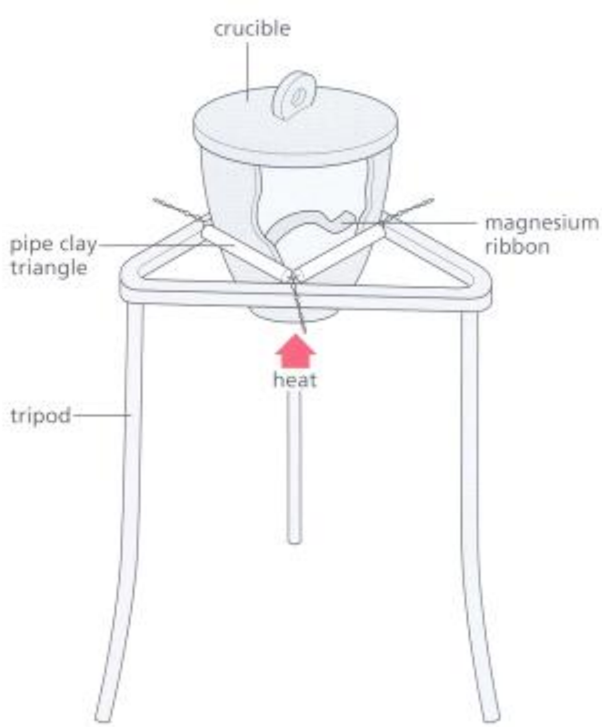


Figure 5.4 Apparatus used to determine magnesium oxide's formula.

Table 5.4 Data from experiment shown in Figure 5.4.

Mass of crucible	14.63 g
Mass of crucible and magnesium	14.87 g
Mass of crucible and magnesium oxide	15.03 g
Mass of magnesium used	0.24 g
Mass of oxygen which has reacted with the magnesium	0.16 g

The data shown in Table 5.1 were obtained from an experiment using the apparatus shown in Figure 5.4 to find the formula for this white powder, magnesium oxide. From these data we can calculate the number of moles of each of the reacting elements. (A_r: O = 16; Mg = 24).

	Mg	**O**
Masses reacting (g)	0.24	0.16
Number of moles	$\frac{0.24}{24}$	$\frac{0.16}{16}$
	= 0.01	= 0.01
Ratio of moles	1	1
Formula	MgO	

This formula is the **empirical formula** of the compound. It shows the simplest ratio of the atoms present.

Unknown compound 1

In another experiment an unknown organic compound was found to contain 0.12 g of carbon and 0.02 g of hydrogen. Calculate the empirical formula of the compound. (A_r: H = 1; C = 12)

	C	**H**
Masses reacting (g)	0.12	0.02
Number of moles	$\frac{0.12}{12}$	$\frac{0.02}{1}$

	= 0.01	= 0.02
Ratio of moles	1	2
Empirical formula	CH_2	

From our knowledge of bonding (Chapter 4) we know that a molecule of this formula cannot exist. However, molecules with the following formulae do exist: C_2H_4, C_3H_6, C_4H_8 and C_5H_{10}. All of these formulae show the same ratio of carbon atoms to hydrogen atoms, CH_2, as our unknown. To find out which of these formulae is the actual formula for the unknown organic compound, we need to know the mass of one mole of the compound.

Using a mass spectrometer, the relative molecular mass (M_r) of this organic compound was found to be 56. We need to find out the number of empirical formulae units present:

M_r of the empirical formula unit

$= (1 \times 12) + (2 \times 1)$

$= 14$

Number of empirical formula units present

$$= \frac{M_r \text{ of compound}}{M_r \text{ of empirical formula unit}}$$

$$= \frac{56}{14}$$

$= 4$

Therefore, the actual formula of the unknown organic compound is $4 \times CH_2 = C_4H_8$.

This substance is called butene. C_4H_8 is the **molecular formula** for this substance and shows the **actual** numbers of atoms of each element

present in one molecule of the substance.

Sometimes the composition of a compound is given as a percentage by mass of the elements present. In cases such as this the procedure shown in the next example is followed.

Unknown compound 2

Calculate the empirical formula of an organic compound containing 92.3% carbon and 7.7% hydrogen by mass. The M_r of the organic compound is 78. What is its molecular formula? (A_r: H = 1; C = 12)

	C	**H**
% by mass	92.3	7.7
in 100 g	92.3 g	7.7 g
moles	$\frac{92.3}{12}$	$\frac{7.7}{1}$
	= 7.7	= 7.7
Ratio of moles	1	1
Empirical formula	CH	

M_r of the empirical formula unit CH

= 12 + 1

= 13

The molecular formula of the organic compound is 6 × CH = C_6H_6. This is a substance called benzene.

Questions

Use the following values of A_r to answer the questions below: H = 1; C =

12; O = 16; Ca = 40.

1. Determine the empirical formula of an oxide of calcium formed when 0.4 g of calcium reacts with 0.16 g of oxygen.

2. Determine the empirical formula of an organic hydrocarbon compound which contains 80% by mass of carbon and 20% by mass of hydrogen. If the M_r of the compound is 30, what is its molecular formula?

Moles and chemical equations

When we write a balanced chemical equation we are indicating the numbers of moles of reactants and products involved in the chemical reaction. Consider the reaction between magnesium and oxygen.

$$\text{magnesium} + \text{oxygen} \rightarrow \text{magnesium oxide}$$

$$2\ Mg\ (s) + O_2\ (g) \rightarrow 2\ MgO(s)$$

This shows that 2moles of magnesium react with 1 mole of oxygen to give 2 moles of magnesium oxide.

Using the ideas of moles and masses we can use this information to calculate the quantities of the different chemicals involved.

$2\ Mg(s) +$	$O_2\ (g)$	$\rightarrow$	$2\ MgO(s)$
2 moles	1 mole		2 moles
2×24	$1 \times (16 \times 2)$		$2 \times (24 + 16)$
= 48 g	= 32 g		= 80 g

You will notice that the total mass of reactants is equal to the total mass of product. This is true for any chemical reaction and it is known as the **Law of conservation of mass**. This law was understood by the Greeks but was first clearly formulated by Antoine Lavoisier in 1774. Chemists can use this idea to calculate masses of products formed and reactants used in chemical processes before they are carried out.

Example using a solid

Lime (calcium oxide, CaO) is used in the manufacture of mortar. It is manufactured in large quantities by a company called Tilcon at their Swindon quarry in North Yorkshire (see Figure 5.5) by heating limestone (calcium carbonate, $CaCO_3$).

Figure 5.5 Calcium oxide production at Swindon quarry.

The equation for the process is:

$CaCO_3\,(s) \rightarrow$	$CaO\,(s)$ +	$CO_2\,(g)$
1 mole	1 mole	1 mole
$(40 + 12 + (3 \times 16))$	$40 + 16$	$12 + (2 \times 16)$
= 100 g	= 56 g	= 44 g

Calculate the amount of lime produced when 10 tonnes of limestone are heated (Figure 5.6). (A_r: C = 12; O = 16; Ca = 40)

1 tonne (t) = 1000 kg

1 kg = 1000g

Figure 5.6 How much lime is produced?

From this relationship between grams and tonnes we can replace the masses in grams by masses in tonnes.

$$CaCO_3\,(s) \rightarrow CaO(s) + CO_2\,(g)$$

	$CaCO_3$	CaO	CO_2
	100t	56t	44t
Hence	10t	5.6t	4.4t

The equation now shows that 100 t of limestone will produce 56 t of lime. Therefore, 10 t of limestone will produce 5.6 t of lime.

Example using a gas

Many chemical processes involve gases. The volume of a gas is measured more easily than its mass. This example shows how chemists work out the volumes of gaseous reactants and products needed using Avogadro's Law and the idea of moles.

Some rockets use hydrogen gas as a fuel. When hydrogen burns in oxygen it forms water vapour. Calculate the volumes of (a) O_2 (g) used and (b) water, H_2O (g), produced if 960 dm^3 of hydrogen gas, H_2 (g), were burned in oxygen. (A_r: H = 1; O = 16) Assume 1 mole of any gas occupies a volume of 24 dm^3.

$2\,H_2$ (g)	$+\,O_2$ (g)	$\rightarrow 2H_2O$ (g)
2 moles	1 mole	2 moles
2×24	1×24	2×24
= 48 dm^3	= 24 dm^3	= 48 dm^3

Therefore:

(× 2)	96 dm^3	48 dm^3	96 dm^3
(× 10)	960 dm^3	480 dm^3	960 dm^3

When 960 dm^3 of hydrogen are burned in oxygen:

a. 480 dm^3 of oxygen are required and

b. 960 dm^3 of H_2O (g) are produced.

Example 1 using a solution

Chemists usually carry out reactions using solutions. If they know the concentration of the solution(s) they are using they can find out the quantities reacting.

Calculate the volume of 1 $mol{\cdot}dm^{-3}$ solution of H_2SO_4 required to react completely with 6 g of magnesium. (A_r: Mg = 24)

number of moles of magnesium

$$= \frac{\text{mass of magnesium}}{\text{mass of 1 mole of magnesium}}$$

$$= \frac{6}{24}$$

= 0.25

Mg(s)	+ H_2SO_4 (aq)	$\rightarrow$ $MgSO_4$ (aq) +	H_2 (g)
1 mole	1 mole	1 mole	1 mole
0.25 mol	0.25 mol	0.25 mol	0.25 mol

We can see that 0.25 mol of H_2SO_4 (aq) is required. Using:

volume of H_2SO_4 (aq) (dm^3)

$$= \frac{\text{moles of } H_2SO_4}{\text{concentration of } H_2SO_4 \ (\text{mol} \cdot \text{dm}^{-3})}$$

$$= \frac{0.25}{1}$$

= 0.25 dm^3 or 250 cm^3

Example 2 using a solution

What is the concentration of sodium hydroxide solution used in the following neutralisation reaction: 40 cm^3 of 0.2 mol·dm^{-3} solution of hydrochloric acid just neutralised 20 cm^3 of sodium hydroxide solution.

number of moles of HCl used

$= \text{concentration (mol·dm}^{-3}) \times \text{volume (dm}^{3})$

$= 0.2 \times 0.04$

$= 0.08$

HCl (aq)	+ NaOH (aq)	$\rightarrow$ NaCl (aq) +	H_2O (l)
1 mole	1 mole	1 mole	1 mole
0.08 mol	0.08 mol	0.08 mol	0.08 mol

You will see that 0.08 mole of NaOH would have been present. The concentration of the NaOH (aq) will be given by:

concentration of NaOH (mol·dm^{-3})

$$= \frac{\text{number of moles of NaOH}}{\text{volume of NaOH (dm}^3)}$$

(volume of NaOH in dm^3 $= \frac{20}{1000} = 0.02$)

$$= \frac{0.08}{0.02}$$

= 4 mol·dm^{-3} or 4 M

Questions

Use the following A_r values to answer the questions below: O = 16; Mg = 24; S = 32; K = 39; Cu = 63.5.

1. Calculate the mass of sulphur dioxide produced by burning 16 g of sulphur in an excess of oxygen in the Contact process.

2. Calculate the mass of sulphur which, when burned in excess oxygen, produces 640 g of sulphur dioxide in the Contact process.

3. Calculate the mass of copper required to produce 159 g of copper (II) oxide when heated in excess oxygen.

4. In the rocket mentioned previously in which hydrogen is used as a fuel, calculate the volume of hydrogen used to produce 24 dm^3 of water ($H_2O(g)$).

5. Calculate the volume of 2 $mol \cdot dm^{-3}$ solution of sulphuric acid required to react with 24 g of magnesium.

6. What is the concentration of potassium hydroxide solution used in the following neutralisation reaction? 20 cm^3 of 0.2 $mol \cdot dm^{-3}$ solution of hydrochloric acid just neutralised 15 cm^3 of potassium hydroxide solution.

Checklist

After studying Chapter 5 you should know and understand the following terms.

Avogadro's Law Equal volumes of all gases measured under the same conditions of temperature and pressure contain equal numbers of molecules.

Calculating moles

Compounds:

$$\frac{\text{mass of compound (in grams)}}{\text{number of moles}} = \text{mass of 1 mole of the compound}$$

$$\text{number of moles} = \frac{\text{mass of compound}}{\text{mass of 1 mole of compound}}$$

Elements:

$$\frac{\text{mass of element (in grams)}}{\text{number of moles}} = \text{mass of 1 mole of the element}$$

$$\text{number of moles} = \frac{\text{mass of the element}}{\text{mass of 1 mole of that element}}$$

Gases:

1 mole of any gas occupies 24 dm^3 (litres) at room temperature and pressure (rtp).

$$\text{number of moles of a gas} = \frac{\text{volume of the gas (in dm}^3\text{ at rtp)}}{24\ \text{dm}^3}$$

Solutions:

$$\text{concentration of a solution (in mol}\cdot\text{dm}^{-3}\text{)} = \frac{\text{number of moles of solute}}{\text{volume (in dm}^3\text{)}}$$

$$\text{number of moles} = \text{concentration (in mol}\cdot\text{dm}^{-3}\text{)} = \text{volume of solution (in dm}^3\text{)}$$

Empirical formula A formula showing the simplest ratio of atoms present.

Mole The amount of substance which contains 6×10^{23} atoms, ions or molecules. This number is called Avogadro's constant.

Atoms - 1 mole of atoms has a mass equal to the relative atomic mass (A_r) in grams.

Molecules - 1 mole of molecules has a mass equal to its relative molecular mass (M_r) in grams.

Molecular formula A formula showing the actual number of atoms of each element present in one molecule.

Relative formula mass (RFM) The sum of the relative atomic masses of all those elements shown in the formula of the substance. This is often referred to as the **relative molecular mass (M_r)**.

Chemical calculations

Additional questions

Use the data in the table below to answer the questions which follow:

Element	**A_r**
H	1
C	12
N	14
O	16
Na	23
Mg	24
Si	28
S	32
Cl	35.5
Fe	56

1. Calculate the mass of:

a. 1 mole of:

(i) chlorine molecules
(ii) iron (III) oxide.

b. 0.5 mole of:

(i) magnesium nitrate
(ii) ammonia.

2. Calculate the volume occupied, at rtp, by the following gases. (One mole of any gas occupies a volume of 24 dm^3 at rtp.)

a. 12.5 moles of sulphur dioxide gas.

b. 0.15 mole of nitrogen gas.

3. Calculate the number of moles of gas present in the following:

a. 36 cm^3 of sulphur dioxide

b. 144 dm^3 of hydrogen sulphide.

4. Using the following experimental information to determine the empirical formula of an oxide of silicon.

Mass of crucible	18.20 g
Mass of crucible + silicon	18.48 g
Mass of crucible + oxide of silicon	18.80 g

5a. Calculate the empirical formula of an organic liquid containing 26.67% of carbon, 2.22% of hydrogen with the rest being oxygen.

b. The M_r of the liquid is 90. What is its molecular formula?

6. Iron is extracted from its ore, haematite, in the blast furnace. The main extraction reaction is:

$$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$$

a. Name the reducing agent in this process.

b. Name the oxide of iron shown in the equation.

c. Explain why this is a **redox** reaction.

d. Calculate the mass of iron which will be produced from 640 tonnes of haematite.

7. Consider the following information about the newly discovered element, vulcium, whose symbol is Vu 'Vulcium is a solid at room temperature. It is easily cut by a penknife to reveal a shiny surface which tarnishes quite rapidly. It reacts violently with water, liberating a flammable gas and forms a solution with a pH of 13. When vulcium reacts with chlorine, it forms a white crystalline solid containing 29.5% chlorine.' (A_r: Vu = 85)

a. Calculate the empirical formula of vulcium chloride.

b. To which group of the periodic table should vulcium be assigned?

c. Write a word and balanced chemical equation for the reaction between vulcium and chlorine.

d. What other information in the description supports the assignment of group you have given to vulcium?

e. What type of bonding is present in vulcium chloride?

f. Write a word and balanced chemical equation for the reaction between vulcium and water.

g. Write the formulae for:

(i) vulcium sulphate
(ii) vulcium carbonate
(iii) vulcium hydroxide.

Look at the periodic table to find out the real name of vulcium.

8. 0.048 g of magnesium was reacted with excess dilute hydrochloric acid at room temperature and pressure. The hydrogen gas given off was collected.

a. Write a word and balanced symbol equation for the reaction taking place.

b. Draw a diagram of an apparatus which could be used to carry out this experiment and collect the hydrogen gas.

c. How many moles of magnesium were used?

d. Using the equation you have written in your answer to a, calculate the number of moles of hydrogen and hence the volume of this gas produced.

e. Calculate the volume of a solution containing 0.1 $mol \cdot dm^{-3}$ hydrochloric acid which would be needed to react exactly with 0.048 g of magnesium.